Chemistry

Key Stage 4

Your name:

Your tutor's name:

Module 1 - Moles

Tutorial	Topic
Tutorial 1.1	Moles
Tutorial 1.2	Amounts of Substances in Equations
Tutorial 1.3	Percentage Yield and Atom Economy
Tutorial 1.4	Moles and Gas Volumes

Learning objectives

In this module you will look at:

1. Understanding and using moles, Avogadro's Constant and mass when calculating the amount of substance.
2. Calculating masses of substances from a balanced equation and vice versa.
3. Calculating and using the percentage yield and atom economy for understanding a reaction pathway.
4. Calculating volumes of gases using molar gas volume (r.t.p) using balanced equations.

Knowledge Check 1

If you'd prefer to complete these questions in an online form, go to:

https://forms.office.com/r/hVEzNBVscK

1. The volume of one mole of any gas at room temperature and pressure is 24.0 dm^3. How many moles of carbon dioxide is 95.0 cm^3?

   A. 3.96×10^{-3} mol
   B. 3.96 mol
   C. 0.253 mol
   D. 253 mol

2. 40 kg of titanium chloride was added to 20 kg of sodium. The equation for the reaction is:
   $TiCl_4 + 4\ Na \rightarrow Ti + 4\ NaCl$
   Relative atomic masses (Ar): Na = 23 Cl = 35.5 Ti = 48

   For a Stage 2 reaction the percentage yield was 92.3%. The theoretical maximum mass of titanium produced in this batch was 13.5 kg. The actual mass of titanium produced was:

   A. 12.5 kg
   B. 12 kg
   C. 13 kg
   D. 13.5 kg

3. A student wanted to make 11.0 g of copper chloride. The equation for the reaction is:
   $CuCO_3 + 2HCl \rightarrow CuCl_2 + H_2O + CO_2$
   Relative atomic masses, Ar: H = 1; C = 12; O = 16; Cl = 35.5; Cu = 63.5

   What is the mass of copper carbonate the student should react with dilute hydrochloric acid to make 11.0 g of copper chloride?

   A. 8.2g
   B. 11 g
   C. 9.7g
   D. 8.7 g

Tutorial 1.1 - Moles

In this tutorial you will look at:

5. Describing the measurement of amounts of substance in moles
6. Calculating number of moles of particles of a substance in a given mass of that substance and vice versa
7. Calculating the number of particles of a substance in a given number of moles or mass of that substance and vice versa

Maths skills for this lesson:

8. Completing calculations with numbers written in standard form

Learning activities

1

Starter Activity

What is a mole and why is it used in chemistry?

One mole of a substance is 6.02×10^{23} of particles of that substance

The type of particles in a substance could be atoms, ions, molecules or electrons.

The Avogadro constant is simply the name given to the number of particles per mole of a substance. It has the value and unit of **6.02×10^{23} mol^{-1}.**

However, one mole of different substances have different masses (in grams) as shown in the table below:

Substance	Number of Moles	Number of particles	Mass of substance (g)
CO_2	1	6.02×10^{23}	44
Na	1	6.02×10^{23}	23
O_2	1	6.02×10^{23}	32

This is because the number of moles of an element is linked to its **relative atomic mass (A_r)** or the number of moles of a molecule is linked to its **relative molecular mass (M_r)** via the following equation:

$$\textbf{Number of Moles} = \frac{\textbf{mass of element or compound (in grams)}}{\textbf{relative atomic mass } (A_r) \textbf{ of element OR relative molecular mass } (M_r) \textbf{ of compound}}$$

Worked Example 1

Calculate the number of moles in 90 g of water

Answer:

Formula of water is H_2O

$M_r = (1 \times 2) + 16 = 18$

$$\text{Number of moles} = \frac{90 \text{ g}}{18} = \underline{5 \text{ mol}}$$

Worked Example 2

Calculate the mass in grams of 3 moles of potassium bromide

Answer:

Formula of potassium bromide is KBr

$M_r = 39 + 80 = 119$

Mass = mol x M_r

= 3 x 119 = <u>357 g</u>

To determine the number of particles (atoms, ions, molecules or electrons) in a substance, the following equation must be applied:

Number of particles = number of moles x Avogadro's constant

Worked Example 3

Calculate the number of molecules in 0.125 mol of carbon dioxide?

Answer:

Number of molecules = number of moles x Avogadro's constant

$= 0.125 \text{ mol} \times 6.02 \times 10^{23} \text{ mol}^{-1} = \underline{7.53 \times 10^{23}}$

2

Guided Example

How many hydrogen atoms are in 0.200 mol of H_2O?

3

Summary Questions

Total Score: …….. /10

4

Exam-Style Questions

1. 6.60 g of carbon dioxide was produced when 3.30 g of methane was burned in sufficient oxygen. Calculate the number of moles of carbon dioxide produced

[2 marks]

2. The value of the Avogadro constant is 6.02×10^{23} mol^{-1}

a) What is meant by the Avogadro constant? **[1 mark]**

b) Calculate the number of water molecules in 72 g of water.

Give your answer to 3 significant figures **[3 marks]**

Total Score: …….. /6

Tutorial 1.2 - Amounts of Substances in Equations

In this tutorial you will look at:

9. Calculating masses of reactants or products from balanced equations, given the mass of one substance
10. Explaining why, in a reaction, the mass of product formed is controlled by the mass of the reactant which is not in excess
11. Deducing the stoichiometry of an equation from the masses of reactants and products

Maths skills for this lesson:

12. Using arithmetic computation and ratio when determining empirical formulae, balancing equations
13. Changing the subject of a mathematical equation

Learning activities

1 **Starter Activity**

1. Balance the equation for this reaction below, in which sodium chlorate(V) may be prepared by passing chlorine into hot aqueous sodium hydroxide.

………Cl_2 + ……..NaOH → ……NaCl + $NaClO_3$ + $3H_2$

2. Balance the equation below, in which nitric acid is formed by the reaction between dinitrogen tetroxide and water.

……..N_2O_4 + ……….H_2O → ………..HNO_3 + 2NO

Once a balanced equation has been created, the coefficients (the numbers in front of each chemical formula) illustrate the amount of each reactant and products **in moles**.

For example, the balanced equation below shows: *2 moles of magnesium metal reacts with 1 mole of oxygen gas to form 1 mole of magnesium oxide*

2Mg(s) + O_2 (g) → MgO (s)

Note: If there is **no coefficient** in front of a chemical formula this means **1 mole** of that substance is present.

<u>Worked Example 1</u>

In a sealed vessel, calcium carbonate decomposes to form calcium oxide and carbon dioxide.

a) How much calcium oxide in grams could be obtained from 800,000 g of calcium carbonate

$CaCO_3$ (s) → CaO (s) + CO_2 (g)

Step 1: Calculate the mass of calcium carbonate that has reacted

Moles of $CaCO_3$ = mass ÷ M_r

= 800,000 g ÷ 100

= 8000 mol

M_r of $CaCO_3$ = A_r(Ca) + A_r(C) + A_r(O x 3)

= 40 + 12 + (16 x 3)

Step 2: Determine the mole ratio of calcium carbonate to calcium oxide by using the coefficients of the balanced equation.

$CaCO_3$ (s) → CaO (s) + CO_2 (g)

1mol : 1 mol

The balanced equation shows mole ratio of $CaCO_3$ to CaO is 1:1

Step 3: Use this information to calculate the theoretical moles of calcium oxide that is produced

$CaCO_3$ (s) → CaO (s) + CO_2 (g)

1 mol : 1 mol

8000 mol : 8000 mol

Therefore when 8000 mol of $CaCO_3$ is used up, 8000 mol of CaO is produced

Step 4: Calculate the mass of calcium oxide produced by using your answer from step 3 and the mole equation

Mass of CaO = moles x M_r of CaO

= 8000 mol x 56

= 448,000 g

M_r of CaO = A_r(Ca) + A_r(O)

= 40 + 16

b) Calculate the mass of CO_2 that is produced

To ensure the law of conservation of mass is applied:

The total mass of reactants = Total mass of products

Mass of CO_2 produced = Mass of $CaCO_3$ – mass of CaO

used up produced

= 800000 g - 448000 g

= 352000 g

In a chemical reaction, the amount of product that is produced is controlled by the limiting reactant. When the limiting reactant gets used up, the reaction stops and no more products can be formed. The number of moles of the limiting reactant can be used to determine the **maximum mass** of product that can be formed.

2

Guided Example

22 g of iron (I) oxide reacts completely with excess carbon monoxide to form iron and carbon dioxide:

$$FeO_2 + 2CO \rightarrow Fe + 2CO_2$$

Calculate the maximum mass of carbon dioxide that can be produced.

Step 1: Determine the number of moles of the limiting reactant

Step 2: Determine the mole ratio of iron (I) oxide to carbon dioxide by using the coefficients of the balanced equation.

Step 3: Use this information to calculate the theoretical moles of carbon dioxide that is produced

Step 4: Calculate the maximum mass of carbon dioxide that is produced by using your answer from step 3 and the mole equation

Worked Example 2

8.1 g of ZnO reacts completely with 0.60 g of C to form 2.2 g of CO_2 and 6.5 g of Zn. Write a balanced symbol equation for this reaction

Answer:

Step 1: Create an unbalanced symbol equation based on the description from the question

…….ZnO +……. C → ……CO_2 + ……. Zn

Step 2: Calculate the moles of all the reactants and products

	…….ZnO +	……. C →	……CO_2	+ ……. Zn
moles = mass/M_r	8.1 g ÷ 81 :	0.60 g ÷ 12 :	2.2 g ÷ 44	: 6.5 g ÷ 65
=	0.10 mol :	0.05 mol :	0.05 mol	: 0.10 mol

Step 3: Divide each mole by the smallest mole value (0.05 mol) to achieve a whole number mole ratio

$$\ldots\ldots ZnO + \ldots\ldots C \rightarrow \ldots\ldots CO_2 + \ldots\ldots Zn$$

$$\text{mol} \div 0.05 = 0.10\text{ mol}\div 0.05 : 0.05\text{ mol}\div 0.05 : 0.05\text{ mol}\div 0.05 : 0.10\text{mol}\div 0.05$$

$$= 2\text{ mol} : 1\text{ mol} : 1\text{ mol} : 2\text{ mol}$$

Step 4: Insert these whole number mole ratios into the balanced equation as coefficients

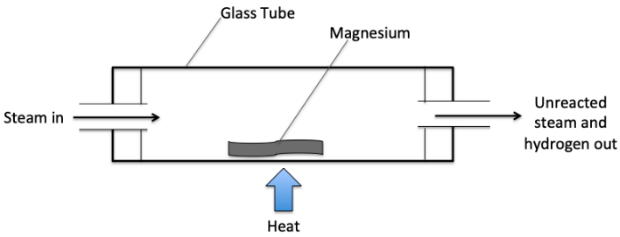
$$\mathbf{2}\ ZnO + C \rightarrow CO_2 + \mathbf{2}\ Zn$$

3 Summary Questions

Total Score: .... /5

4 Exam-Style Questions

1. Magnesium and steam react to produce magnesium oxide and hydrogen gas. A student carried out the reaction in the apparatus shown.



a) Deduce a balanced symbol equation for this reaction including state symbols **[2 marks]**

b) The student used 2.00 g of magnesium in the reaction. Calculate the maximum mass of magnesium oxide produced. Give your answer to 3 significant figures (s.f) **[3 marks]**

2. Citric acid ($C_6H_8O_7$) and sodium hydrogencarbonate ($NaHCO_3$) react as shown in the equation: $C_6H_8O_7$ (s) + $3NaHCO_3$ (s) → $3CO_2$ (g) +$3H_2O$ (l) + $C_6H_5O_7Na_3$(aq)

a) Calculate the minimum mass of sodium hydrogencarbonate that is needed to ensure that all of 20.0 g of citric acid used up in the reaction. Give your answer to 3 s.f. **[4 marks]**

b) In a second repeat of this reaction a different mass of citric acid was used. At the end of this second repeat 7.2 g of water was collected. Calculate the mass of citric acid that reacted. Give your answer to 3 significant figures **[4 marks]**

3. The catalyst AO_x has an M_r of 79.5 and can be made by reacting Metal A with oxygen gas. The reaction is carried out using 14.10 g of Metal A in an excess of oxygen gas. At the end of the reaction 17.65 g of AO_x had formed with no other products.

a) Determine the formula of the catalyst (AO_x) and metal A **[6 marks]**

b) Use your answer to part i) to write a balanced symbol equation for this reaction including state symbols **[2 marks]**

Total Score: …….. /21

Tutorial 1.3 - Percentage Yield and Atom Economy

In this tutorial you will look at:

14. Calculating the percentage yield of a reaction from the actual yield and the theoretical yield
15. Explaining why the actual yield of a reaction is usually less than the theoretical yield
16. Calculating the atom economy of a reaction to form a desired product from the balanced equation
17. Explaining why a particular reaction pathway is chosen to produce a specified product given appropriate data

Maths skills for this lesson:

18. Completing arithmetic computation when calculating yields and atom economy

Learning activities

1

Starter Activity

1. What does the term 'yield' mean?

2. What does it mean if a reaction has 100% percentage yield?

3. What do we need to know before we can work out the percentage yield of a reaction?

The percentage yield of a reaction tells us about the overall success of an experiment. It is a percentage that compares the maximum **theoretical yield** and **the actual yield. The theoretical yield** = the maximum possible mass of a product that can be made in a chemical reaction. This can be calculated from the balanced symbol equation (see tutorial 2). **The actual yield** = the actual mass of the product you obtain from the reaction.

The percentage yield of a reaction can be worked out via the following equation:

$$\textbf{Percentage Yield} = \frac{\textbf{Actual Yield}}{\textbf{Theoretical Yield}} \textbf{x100}$$

Another way of writing the same equation is:

$$\textbf{Percentage Yield} = \frac{\textbf{Mass of actual product made}}{\textbf{Maximum theoretical mass of product}} \textbf{ x100}$$

Worked Example 1

A company making sulfuric acid gets an actual yield of 74 tonnes. They predicted a yield of 85 tonnes. What is the percentage yield of this reaction?

Answer:

$$\textbf{Percentage Yield} = \frac{\textbf{Actual Yield}}{\textbf{Theoretical Yield}} \textbf{ x100} \quad = \frac{\textbf{74 tonnes}}{\textbf{85 tonnes}} \textbf{ x100}$$

$$= 87\%$$

Worked Example 2

A chemist used 100g of salicylic acid to synthesise aspirin via the following chemical reaction:

$$C_7H_6O_3 + C_4H_6O_3 \rightarrow C_9H_8O_4 + CH_3COOH$$

salicylic acid — aspirin

When the chemist completed the reaction the actual mass of aspirin they obtained was 85 g. Calculate the percentage yield of aspirin. Give your answer to 2 significant figures.

The M_r of salicylic acid ($C_7H_6O_3$) is 138.
The M_r of aspirin ($C_9H_8O_4$) is 180.

Answer:

Step 1: Calculate the theoretical yield of aspirin/maximum theoretical mass of aspirin.

Moles of $C_7H_6O_3$ = 100 g ÷138 = 0.725 mol

Mole ratio of $C_7H_6O_3$ to $C_9H_8O_4$ is 1:1

Therefore 0.725 mol of $C_9H_8O_4$ is expected to be made

Maximum theoretical mass of aspirin = mol x M_r

= 0.725 mol x 180

= 130.5 g

Step 2: Determine the actual yield/mass of product actually made from the question:

Actual yield/mass of product actually made = 85 g

Step 3: Calculate the percentage yield of the reaction

$$\text{Percentage Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100 = \frac{85g}{130.5} \times 100$$

$$= \underline{65\%}$$

The reasons why percentage yields are less than 100% are:

- The reaction doesn't go to completion because the reaction is reversible
- Some of the product may be lost when it is separated from the reaction mixture e.g. lost during filtration
- Some of the reactants may react in ways different to the expected reaction i.e. Unexpected side reactions may occur

Atoms cannot be created or destroyed as highlighted by the law of conservation of mass. However not all atoms from reactants will end up in the desired product. In industry, chemists are interested in obtaining a high percentage atom economy of a reaction, which means ensuring a majority of the atoms in reactants end up in the desired product.

The percentage atom economy tells us the percentage of atoms in the total reactants that make up the desired product. The percentage atom economy of a reaction can be worked out via the following equation:

$$\textbf{Percentage Atom Economy} = \frac{\textbf{total } M_r \textbf{ of the desired product}}{\textbf{total } M_r \textbf{ of all the reactants}} \times 100$$

Note: Percentage atom economy can also referred to as 'atom economy'

The highest possible value of atom economy is 100%, when all the reactant atoms end up in the desired product

2 **Guided Example**

Calculate the percentage atom economy to produce iron:

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

High percentage atom economies are important because they have economic, ethical and environmental benefits for industry and society, such as:

- Less waste is produced
- Manufacturing costs are reduced

By comparing the percentage atom economies of different reactions that produce the same desired product, industrial chemists can make informed decisions about which reaction method they should use on a larger scale.

3

Guided Example

A company has 2 methods of making magnesium sulfate

Method 1: $MgO + H_2SO_4 \rightarrow MgSO_4 + H_2O$

Method 2: $MgCO_3 + H_2SO_4 \rightarrow MgSO_4 + H_2O + CO_2$

Which method should the company select to mass produce magnesium sulfate? Show your working out.

4

Summary Questions

Total Score: …….. /10

5

Exam-Style Questions

1. A student aims to synthesise hydrated calcium sulfate crystals via the reaction method shown below

Step 1: Thermal decompose calcium carbonate

$CaCO_3$ (s) → CaO (s) +CO_2 (g)

Step 2: Neutralise calcium oxide with an excess of sulfuric acid

CaO (s) + H_2SO_4 (aq) → $CaSO_4$ (aq) + H_2O (l)

Step 3: Filter the solution, then crystalise the filtrate by heating the water.

a) For **step 1**, the student starts off with 24.00 g of calcium carbonate. At the end of step 1 the student obtained 4.75g of calcium oxide. Calculate the percentage yield of calcium oxide. Give your answer to 3 significant figures. **[6 marks]**

b) Calculate the percentage atom economy for step 2. Give your answer to 1 decimal place **[3 marks]**

c) Another reaction method to synthesis calcium sulfate involves the following steps:
Step 1: Add calcium carbonate to sulfuric acid until all the acid has reacted.
Step 2: Filter the solution, then follow it to crystallise at room temperature
The student says the 3-step process would give a higher percentage yield than the 2-step process. Explain whether or not the student is correct. **[2 marks]**

2. a) Evaluate which process you would select for the mass production of ethanoic acid **[3 marks]**

1. Oxidation of Butane:

$2C_4H_{10}\,(g) + 5O_2\,(g) \rightarrow 4CH_3COOH(l) + 2H_2O\,(l)$

2. Ester Hydrolysis

$CH_3COOCH_2CH_2CH_3\,(l) + H_2O(l) \xrightarrow{\text{Heat}} CH_3COOH\,(l) + CH_3CHCH_2OH\,(l)$

b) A chemist has decided to go with the ester hydrolysis method for the mass production of ethanoic acid. Suggest one thing the chemist could do to help reduce the cost associated with this method **[1 mark]**

Total Score: …….. /15

Tutorial 1.4 - Moles and Gas Volumes

In this tutorial you will look at:

19. Describing the relationship between molar amounts of gases and their volumes and vice versa
20. Calculating the volumes of gases involved in reactions using the molar gas volume at room temperature and pressure
21. Calculating volumes of gases from a balanced equation and a given volume or mass of a reactant or product

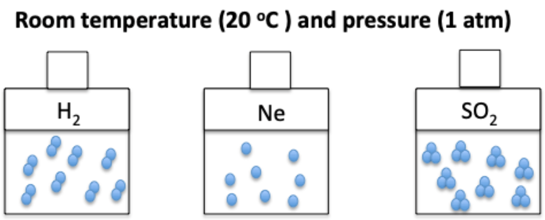
Maths skills for this lesson:

22. Converting units where appropriate particularly from mass to moles
23. Completing arithmetic computation, ratio, percentage and multistep calculations permeates quantitative chemistry
24. Providing answers to an appropriate number of significant figures

Learning activities

1

Starter Activity



Three gases, H_2, Ne and SO_2 occupy the same volume of 24 dm^3 at room temperature and pressure (r.t.p)

a) Sally says 'Since all three gases occupy the same volume at r.t.p., they must all have the same mass'. Is Sally correct? Explain your answer.

b) Jill says 'since all three gases occupy the same volume at RTP, they must all have the same number of moles'. Is Jill correct? Explain your answer.

At room temperature and pressure, Avogadro's law specifically states:

The volume of one mole of any gas is 24 dm^3

Note: Room temperature and pressure (r.t.p.) = 20 °C and 1 atmosphere pressure (1 atm). Therefore at r.t.p. one mole of gas has a molar gas volume of 24 dm^3 mol^{-1} or 24000 cm^3 mol^{-1}. This can be illustrated by the following equation:

$$\textbf{Number of moles of gas} = \frac{24\ dm^3}{24\ dm^3 mol^{-1}} = 1\ mol$$

The following equation is true for all gases at r.t.p:

$$\textbf{Number of moles of gas} = \frac{\text{Volume } (dm^3)}{24\ dm^3 mol^{-1}} = \frac{\text{Volume } (cm^3)}{24000\ cm^3 mol^{-1}}$$

Thus, we can calculate the number of moles of any gas at r.t.p provided its volume is known and vice versa.

At different temperature and pressures the molar gas volume can change. However, Avogadro's Law states that at any particular temperature and pressure equal volumes of different gases contain an equal number of molecules. Therefore, the following generic equation can be applied for different temperatures and pressures.

$$\textbf{Number of moles of gas} = \frac{\text{Volume } (dm^3)}{\text{Molar gas volume } (dm^3 mol^{-1})}$$

Worked Example 1:

Calculate the volume of 0.20 mol of oxygen at room temperature and pressure

Answer:

Volume (dm^3) = number of moles of gas x 24 dm^3 mol^{-1}

= 0.20 mol x 24 dm^3 mol^{-1} = 4.8 dm^3

Worked Example 2

Calculate the number of moles of carbon dioxide gas that occupies 360 cm^3 at room temperature and pressure.

Answer:

$$\text{Number of moles of gas} = \frac{\textbf{Volume } (cm^3)}{\mathbf{24000\ cm^3 mol^{-1}}}$$

$$= \frac{360\ (cm^3)}{24000\ cm^3 mol^{-1}}$$

= 0.015 mol

NOTE:

There is no need to include the units when showing your working out.

E.g. writing the following is acceptable:

$$\textbf{Number of moles} = \frac{360}{24000}$$

= 0.015 mol

OR

360/1000 = 0.36 dm^3

Number of moles of gas = $\frac{\textbf{Volume (dm}^3\textbf{)}}{\textbf{24 dm}^3\textbf{mol}^{-1}}$

$$= \frac{0.36 \text{ dm}^3}{24 \text{ dm}^3\text{mol}^{-1}} = \underline{0.015 \text{ mol}}$$

Remember: a balanced symbol equation can help us to determine the moles of a product or reactant based on the mass of another substance the has been provided.

Similarly, molar gas volumes can be used to calculate the volumes of gases from a balanced equation, given the volume or mass of a reactant or product has been provided

2

Guided Example

4.6 g of sodium reacts completely with water to form aqueous sodium hydroxide and hydrogen gas

$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

Calculate the volume in dm^3 of hydrogen produced (molar volume = 24 dm^3)

Worked Example 3

This reaction takes place in a car's catalytic converter:

$$2CO\ (g) + 2NO\ (g) \rightarrow 2CO_2\ (g) + N_2\ (g)$$

What is the total volume of gases produced when 12 dm^3 of carbon monoxide reacts at r.t.p?

Answer:

Important: Since no masses have been provided to calculate moles using mol = mass/M_r , we must work directly with the mole ratios of the balanced equation to determine the total volume of gases produced.

- The mole ratio of $CO:CO_2$ is 1:1
- Therefore if 12 dm^3 of CO reacts then **12 dm^3 of CO_2 is produced**
- The mole ratio of CO: N_2 is 2:1
- Therefore if 12 dm^3 of CO reacts then **6 dm^3 of N_2 is produced**

Therefore, the total volume of gas that is produced is

- **6 dm^3 + 12 dm^3 = <u>18 dm^3</u>**

3

Summary Questions

Total Score: …….. /11

4

Exam-Style Questions

1. The following acid base reaction is carried out at room temperature and pressure:

$$CaCO_3\,(s) + H_2SO_4\,(aq) \rightarrow CaSO_4\,(aq) + H_2O\,(l) + CO_2\,(g)$$

0.400g of calcium carbonate is reacted with an excess of sulfuric acid. Calculate the volume in dm^3 of carbon dioxide gas that is produced (The volume of one mole of gas at room temperature and pressure is 24.0 dm^3) **[3 marks]**

2. a) The decomposition of hydrogen peroxide can be illustrated by the following equation:

$$2H_2O_2\ (aq) \rightarrow 2H_2O\ (l) + O_2\ (g)$$

When this decomposition reaction is carried out at room temperature and pressure, 60 cm^3 of oxygen gas is produced. Calculate the number of moles of oxygen gas produced (The volume of one mole of gas at room temperature and pressure is 24.0 dm^3) **[2 marks]**

b) Calculate the maximum volume of oxygen gas that could be obtained from 48 g of hydrogen peroxide (The volume of one mole of gas at room temperature and pressure is 24.0 dm^3) **[3 marks]**

3. Ammonia, hydrogen and nitrogen gas all have different boiling points, but at the same temperature and pressure it was found that 1 mole of each gas occupied the same volume. Explain why. **[3 marks]**

4. The synthesis of ammonia takes place via the Haber process. This is illustrated by the following equation:

$$N_2\ (g) + 3H_2\ (g) \rightleftharpoons 2NH_3\ (g)$$

Calculate the volume in cm^3 of ammonia produced from the reaction of 795 dm^3 of hydrogen (The volume of one mole of gas at room temperature and pressure is 24.0 dm^3). **[3 marks]**

Total Score: …….. /14

Module 1 review

How do you feel now?

In this module we looked at:

1. Understanding and using moles, Avogadro's Constant and mass when calculating the amount of substance.
2. Calculating masses of substances from a balanced equation and vice versa.
3. Calculating and using the percentage yield and atom economy for understanding a reaction pathway.
4. Calculating volumes of gases using molar gas volume (r.t.p) using balanced equations.

Reflection is important because it helps you review and improve the way you approach tasks, rather than just carrying on doing things as you have always done them.

Take a few minutes to think about where you think you did well during the last four tutorials, and what you think you could improve on. Write you thoughts in the boxes below.

What I did well…	What I could have improved on…

What I could do differently to make the most of my next tutorials…

Knowledge Check 2

If you'd prefer to complete these questions via an online form, go to:

https://forms.office.com/r/QrFWxSdvi4

1. In industry, methanol is produced by reacting carbon monoxide with hydrogen. The equation for the reaction is:
   $CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$
   How many moles of carbon monoxide react completely with 4.0×10^3 moles of hydrogen?

   A. 1.0×10^3 moles
   B. 2.0×10^3 moles
   C. 4.0×10^3 moles
   D. 8.0×10^3 moles

2. The equation for the complete combustion of ethane is:
   $2C_2H_6\,(g) + 7O_2\,(g) \rightarrow 4CO_2\,(g) + 6H_2O(l)$
   Which is the correct volumes of ethane and oxygen that react together and the volume of carbon dioxide they produce when they react as shown in this equation? (all volumes of gases are measured under the same conditions of temperature and pressure)

   A. Ethane: 5, Oxygen: 35, Carbon dioxide: 10
   B. Ethane: 5, Oxygen: 70, Carbon dioxide: 20
   C. Ethane: 10, Oxygen: 35, Carbon dioxide: 20
   D. Ethane: 10, Oxygen: 70, Carbon dioxide: 40

3. Look at the equations for the two reactions:
   Reaction 1: $CuCO_3(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + H_2O(l) + CO_2(g)$
   Reaction 2: $CuO(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + H_2O(l)$
   What is the percentage atom economy for Reaction 2?

   A. 89.20%

   B. 85.20%

   C. 88.20%

   D. 87.20%

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